

# *Chemistry 1B*

*Fall 2016*

Topic 23

[more] Chemical Kinetics



- kinetics and mechanism of chemical reaction
- energy profile and reaction coordinate
- activation energy and temperature dependence of rate constant
- catalysis

# ***chemical kinetics***

***elementary reactions***

***reaction mechanisms***

***reaction energy profile and the reaction  
coordinate diagram***

## Learning Objectives and Worksheet XII

### Chemistry 1B-AL Fall 2016

#### Lectures 21-22) Chemical Kinetics

To date CHEM1B-AL has focused on the structure and chemical properties of atoms and molecules. Chemical kinetics, the final topic of the course, shifts attention to chemical reactions. In particular we will study how measuring the factors affecting the rate of a chemical reaction gives clues to the series of individual steps, i.e. the mechanism, by which the reaction proceeds.

#### I. Basic Concepts and Definitions

- For a chemical reaction, what are the differences between the information that is provided by thermodynamics [eq *at equilibrium*] (CHEM-1A, CHEM-1C) and kinetics?

- In a kinetics experiment one measures decreases in \_\_\_\_\_ or increases in \_\_\_\_\_ vs \_\_\_\_\_.

#### II. Differential rate laws

- The rate at which a reaction occurs is written in terms of the \_\_\_\_\_ of the \_\_\_\_\_ reactants (raised to powers) since for higher \_\_\_\_\_ there are a greater number of molecular \_\_\_\_\_ allowing the possibility that the reaction will proceed more rapidly.

- In the differential rate expression:

$$-\frac{d[X]}{dt} = +n\frac{d[Y]}{dt} = k_f[A]^m[B]^n - k_r[C]^k[D]^l$$

- $-\frac{d[X]}{dt}$  represents the \_\_\_\_\_ of a \_\_\_\_\_ X.

- $\frac{d[Y]}{dt}$  represents the \_\_\_\_\_ of a \_\_\_\_\_ Y.

- $k_f$  and  $k_r$  are the \_\_\_\_\_ and \_\_\_\_\_ respectively.

- the order of the above reaction in reactant A is \_\_\_\_\_.

- The expression

$$-\frac{d[A]}{dt} = k_f[A]_0^m[B]_0^n$$

describes the rate of reaction when \_\_\_\_\_ and thus is called the \_\_\_\_\_ rate law. Here the total rate order is \_\_\_\_\_.

- If  $-\frac{d[A]}{dt} = k_f[A]^m$ , the instantaneous rate of reaction is given by the \_\_\_\_\_ of a plot of [A] vs t.

#### III. Determining order of a reaction

- For a reaction all reactant concentrations are kept constant except for  $[A]_0$ .
  - If the initial rate increases 9-fold when  $[A]_0$  is tripled (i.e.  $[A]_0$  increased 3-fold), the order of the reaction in [A] is \_\_\_\_\_.
  - If the reaction was third-order in [A] and the concentration of  $[A]_0$  is doubled, the initial rate of reaction increases by a factor of \_\_\_\_\_.
  - If the rate of reaction doesn't change when  $[A]_0$  is doubled, the order of the reaction in reactant A is \_\_\_\_\_.
- A reaction has an initial reaction rate (loss of [A]) that is first-order in  $[A]_0$  and second-order in  $[B]_0$ . For the initial concentrations of  $[A]_0=1\text{ M}$  and  $[B]_0=2\text{ M}$  the initial rate is  $\frac{20\text{ mol}}{\text{L sec}}$ . The rate constant for the reaction is \_\_\_\_\_.

#### IV. Integrated rate expressions

- While the **differential** rate expression gives the \_\_\_\_\_ of the concentration of a reactant or product with \_\_\_\_\_, the **integrated** rate expression gives the total \_\_\_\_\_ of the reactant or product as a function of \_\_\_\_\_.



2. Since integral calculus is not a prerequisite for CHEM1B (however you better take this important class!), you will be given the formulas for the integrated rate expressions will be supplied on your exams. However you should recognize:

For the reaction where effectively only reactant [A] varies:

$$-\frac{d[A]}{dt} = k[A]^x$$

- i. for  $\ln[A] = -kt + \ln[A]_0$  the order  $x =$  \_\_\_\_\_
  - ii. for  $\frac{1}{[A]} = kt + \frac{1}{[A]_0}$  the order  $x =$  \_\_\_\_\_
  - iii. for  $[A] = -kt + [A]_0$  the order  $x =$  \_\_\_\_\_
3. For the integrated rate expression:
- i. for  $\ln[A] = -kt + \ln[A]_0$  a plot of  vs \_\_\_\_\_ would be a straight line with slope \_\_\_\_\_.
  - ii. for  $\frac{1}{[A]} = kt + \frac{1}{[A]_0}$  a plot of  vs \_\_\_\_\_ would be a straight line with slope  $k$  \_\_\_\_\_.
  - iii. for  $[A] = -kt + [A]_0$  a plot of  vs \_\_\_\_\_ would be a straight line with slope \_\_\_\_\_.

#### V. Kinetics and the mechanism of a chemical reaction

In O-Chem and beyond you will be studying how different kinetic analyses are used to determine a variety of reaction mechanisms. In CHEM1B we will get a 'taste' of the process of kinetics → mechanism for a multi-step reaction but limit the playing field to one type of process- a slow initial step followed by a fast reaction.

#### VI. Elementary reactions

1. An elementary reaction is one that corresponds to \_\_\_\_\_.
2. When the measured rate law for a chemical reaction does not correspond to the reaction stoichiometry the reaction is \_\_\_\_\_ an elementary reaction. When the measured rate law for a chemical reaction corresponds to the reaction stoichiometry the reaction is \_\_\_\_\_ an elementary reaction.

3. An elementary reaction involving only the interaction (collision) between two reactants is \_\_\_\_\_ and has a \_\_\_\_\_ order overall rate.
4. An elementary reaction involving only decomposition of a single reactant \_\_\_\_\_ and has a \_\_\_\_\_ order overall rate.
5. An elementary reaction that requires the improbable collision of three molecular species is a \_\_\_\_\_ and is a \_\_\_\_\_ elementary reaction.

#### VII. Temperature dependence of the rate constant

1. The Arrhenius equation is  $k = Ape^{-E_a/RT}$
2. The factor  $A$  is related to the \_\_\_\_\_ of the molecules participating in the reaction.
3. The magnitude of factor  $A$  depends on how the reaction depends on the \_\_\_\_\_ of the molecules interacting in the reaction.
4. The term  $e^{-E_a/RT}$  enters into the rate giving the relative \_\_\_\_\_.
5. At a given temperature a larger activation energy  $E_a$  will result in \_\_\_\_\_ collisions having \_\_\_\_\_.
6. For a given activation energy  $E_a$  a higher temperature  $T$  will result in \_\_\_\_\_ collisions having \_\_\_\_\_.
7. A maxima in the reaction profile (energy vs extent of reaction, reactants → products), corresponds to the \_\_\_\_\_ for the reaction.
8. The energy difference between the reactants and the maximum is the \_\_\_\_\_.
9. The energy difference between the products and the maximum is the \_\_\_\_\_.
10. A relative minimum in the reaction profile is a \_\_\_\_\_.

#### VIII. Catalysis

1. In general, a catalyst increases the rate of reaction by allowing a reaction pathway that \_\_\_\_\_.
2. What are three types of catalysis
  - i.
  - ii.
  - iii.

- elementary reactions
- kinetics and the mechanism of a chemical reaction
- energy profile and reaction coordinate



*if the reaction was a collision between a NO<sub>2</sub> molecule and a CO molecule one might expect for the differential rate law:*

$$-\frac{d[\text{NO}_2]}{dt} = k[\text{NO}_2][\text{CO}] \quad \times$$

but the observed rate of reaction is:

$$-\frac{d[\text{NO}_2]}{dt} = k[\text{NO}_2]^2$$

**You might ask ?**



the observed kinetics  $-\frac{d[NO_2]}{dt} = k[NO_2]^2$



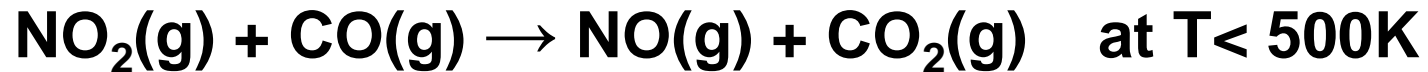
is not consistent with this stoichiometric equation in representing the actual molecular collisions involved in the reaction

ok?? but first let's look at reactions where the stoichiometric equation **DOES** represent the actual collision process. these are called **ELEMENTARY REACTIONS**





back to the question: so what's up DOC ??



answer: the above (stoichiometric) equation  $\text{NO}_2$  colliding with  $\text{CO}$  does NOT represent the molecular mechanism (molecular steps of the reaction)

in actuality: the reaction takes place in two steps

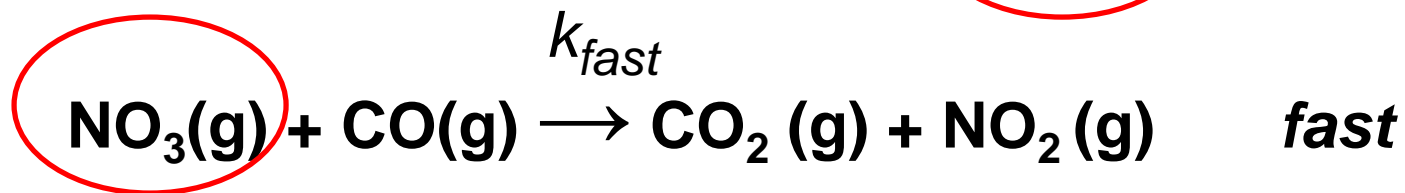
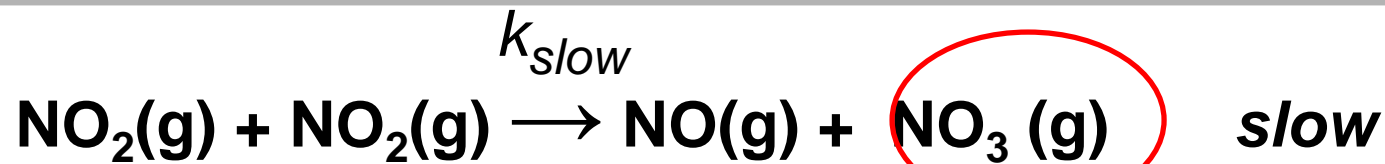


**net  
reaction**



*mechanism of the reaction*  $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{CO}_2(\text{g})$  *overall reaction* ( $T < 500\text{K}$ )

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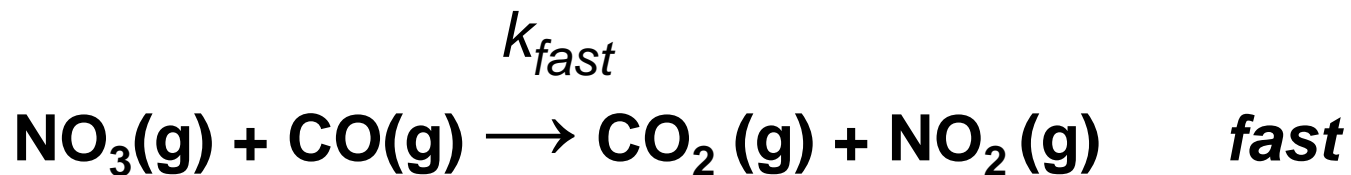


- The two steps are the actual molecular processes by which the reaction occurs and is the **MECHANISM** the reaction
- They represent **ELEMENTARY** reactions
- The combination of elementary reactions must sum to the overall stoichiometry
- Species [e.g.  $\text{NO}_3(\text{g})$ ] appearing in the steps of the mechanism, but not the overall reaction are **REACTION INTERMEDIATES**

## how to get rate law from mechanism of the reaction

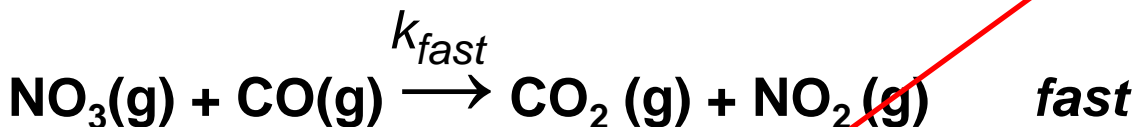
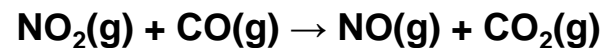
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T < 500K



- rates for individual steps come directly from the stoichiometry of the **ELEMENTARY** reactions
- how to combine rates of elementary reactions to get overall rate equation  
LOTS OF TRICKS (but we don't have time for many)
- our **ONE** example will be **slow** reaction followed by **rapid** reaction (like above reaction)

## how to get rate law from mechanism of the reaction



since these are ELEMENTARY reactions:

$$-\frac{d[\text{NO}_2]}{dt} = k_{\text{slow}} [\text{NO}_2]^2 = +2 \frac{d[\text{NO}_3]}{dt} \quad \text{slow}$$

$$-\frac{d[\text{CO}]}{dt} = k_{\text{fast}} [\text{NO}_3][\text{CO}] \quad \text{fast}$$

- since  $k_{\text{fast}} \gg k_{\text{slow}}$  as soon as slow reaction produces  $\text{NO}_3$  the fast reaction occurs
- the total rate is the just the rate of the slow step  
total rate =  $k_{\text{slow}} [\text{NO}_2]^2$  (2<sup>nd</sup> order in  $\text{NO}_2$ , 0<sup>th</sup> order in  $\text{CO}$ )

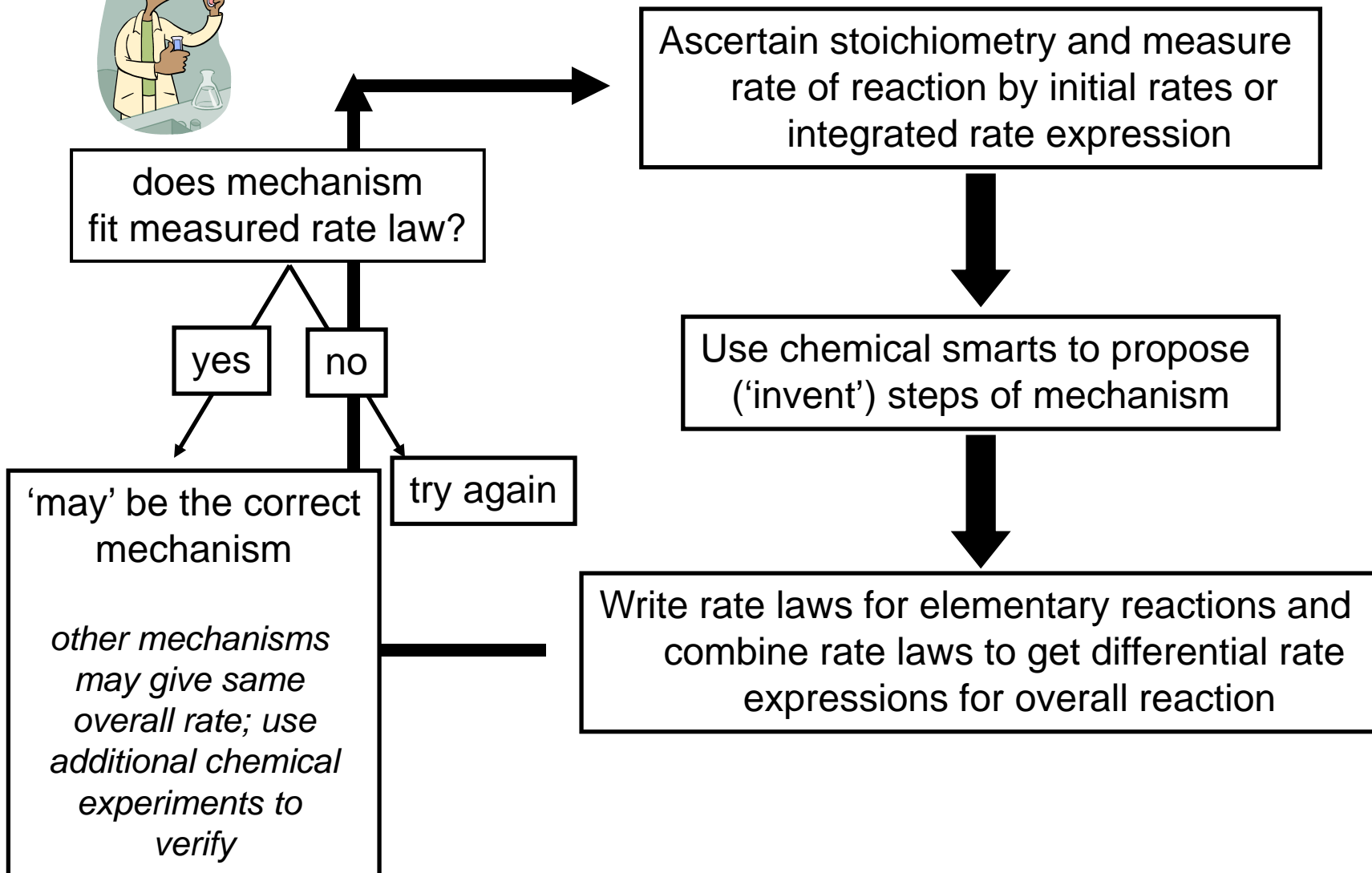
from slide #5: the observed rate of reaction is:  $-\frac{d[\text{NO}_2]}{dt} = k[\text{NO}_2]^2$



- in general the rate law cannot be written in terms of the stoichiometric coefficients of the overall reaction
- the rate law for an elementary reaction **IS** determined by the molecularity (stoichiometry) of the individual step
- the overall rate law is obtained from combining rate laws of elementary reactions
- an overall reaction where the measured rate law and the stoichiometric coefficients 'match' **MAY** or **MAY NOT** be an elementary reaction

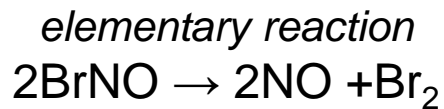


# *the actual process of determining mechanism of a reaction*

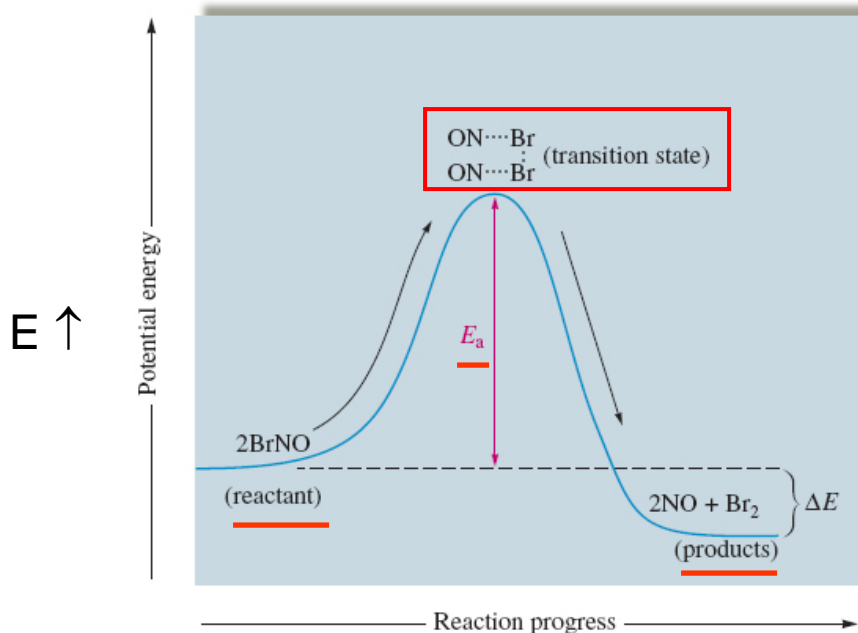




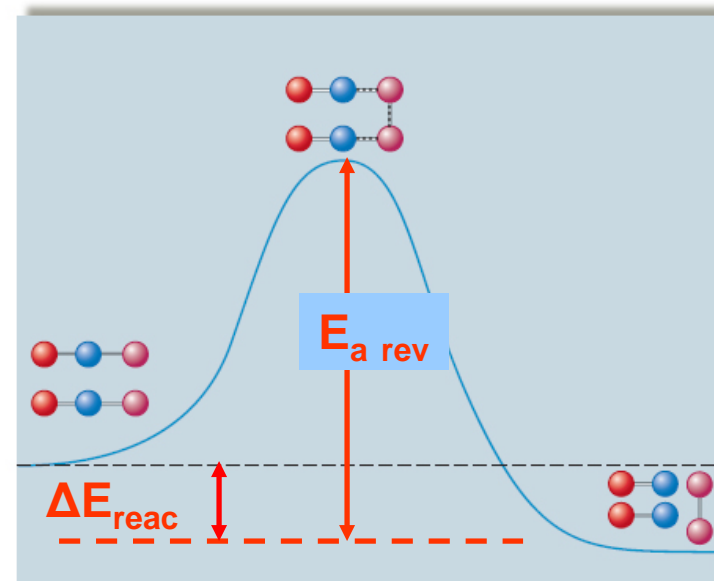
# the reaction coordinate and course of a chemical reaction



molecularity?  
 bimolecular



(a)



(b)

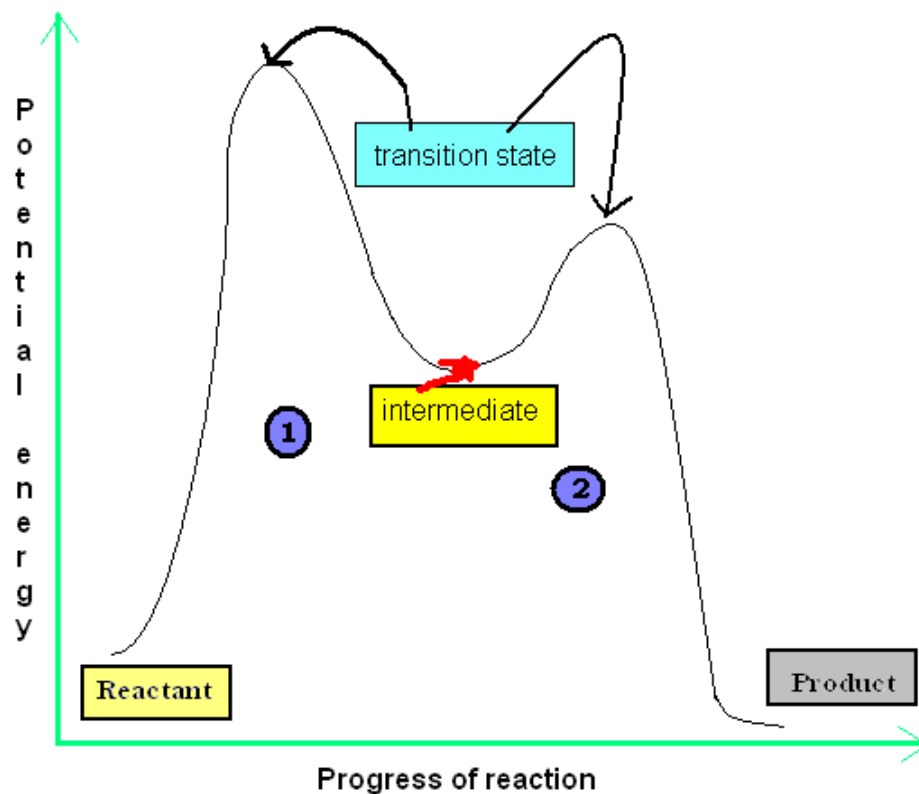
- **energy vs reaction coordinate** (progress of reaction reactants –products)
- **transition state** (maxima of E vs reaction coordinate)

*[N---Br bonds breaking, Br---Br bond forming]*

- **activation energy** ( $E_a = E_{\text{transition state}} - E_{\text{reactants}}$ )
- **reverse activation energy** ( $E_{a \text{ rev}} = E_{\text{transition state}} - E_{\text{products}}$ )
- **$\Delta E_{\text{reaction}}$  ( $\Delta H$ )** ( $= E_{\text{products}} - E_{\text{reactants}}$ )

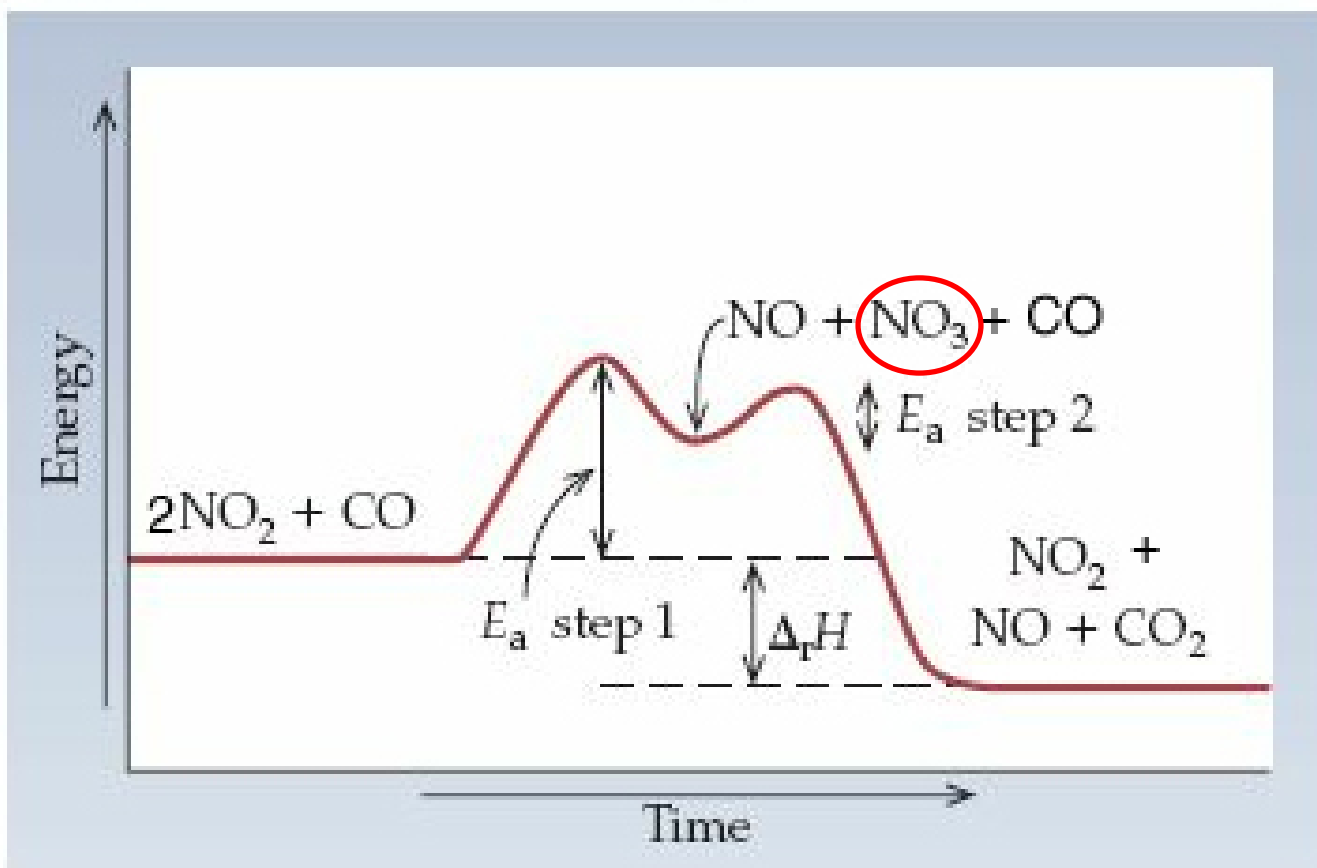


## multi-step reaction



(‘stable’) **Reaction Intermediate** is at a relative minimum along curve

<http://chemwiki.ucdavis.edu/@api/deki/files/916/=intermediate.bmp>

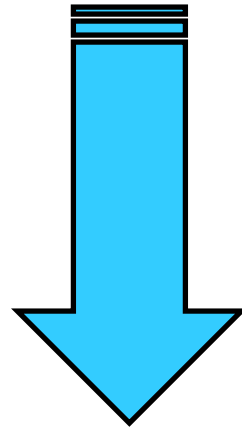




**our  
last  
video !!**

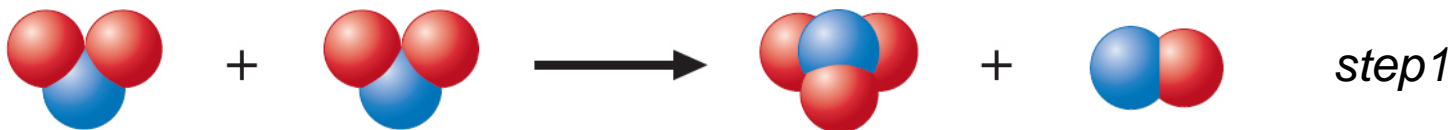


***onto more kinetics !!!***

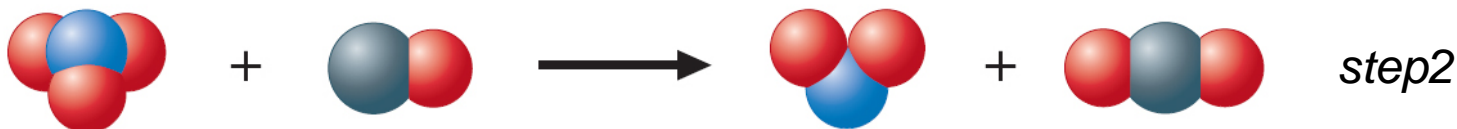




Step 1



Step 2



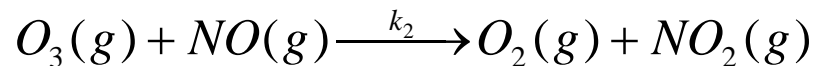
- an elementary reaction represents the molecules ‘colliding’ in an actual step of the overall reaction
- the rate law for an elementary reaction **is** determined by the stoichiometry of the elementary reaction (the order of each reactant is its stoichiometric coefficient)
- the overall order of an elementary reaction is its **molecularity**
- the ‘sum’ of the individual elementary reactions must yield the equation for the overall reaction

## elementary reactions and 'molecularity'

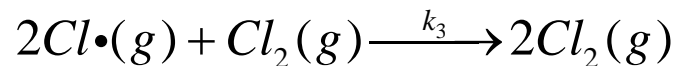
elementary reaction



$$-\frac{d[{}^{14}_6\text{C}]}{dt} = k_1 [{}^{14}_6\text{C}] \quad \text{first-order}$$



$$-\frac{d[\text{O}_3(\text{g})]}{dt} = k_2 [\text{O}_3][\text{NO}] \quad \text{second-order}$$



$$-\frac{d[\text{Cl}_2(\text{g})]}{dt} = k_3 [\text{Cl}\cdot]^2 [\text{Cl}_2] \quad \text{third-order}$$

molecularity

**unimolecular**

**bimolecular**

**termolecular**

*improbable*



if  $A + 2B \rightarrow C$  **IS** an elementary reaction  
then

$$-\frac{d[A]}{dt} = k[A][B]^2 \quad \text{termolecular} \quad \boxed{\text{TRUE}}$$

---

if the measured rate law for  $A + 2B \rightarrow C$  is

$$-\frac{d[A]}{dt} = k[A][B]^2 \quad \text{third order}$$

then

$A + 2B \rightarrow C$  **MAY** or **MAY NOT** be **TRUE**  
an elementary reaction



*a multi-step mechanism for reaction may have a set of elementary reactions that coincidentally leads to the third-order rate law*





- Temperature dependence of rate constant
- Catalysis

# collision theory and effect of temperature on rate constant

---

$$k = zp e^{-E_a/RT}$$

collision  
frequency  
*(collisions per sec)*

orientation factor  
*will colliding molecules be  
in the right orientation to react ?*

probability that a collision will have  
energy  $> E_a$   
*(will get over activation barrier at  
transition state and become products)*



# Arrhenius equation

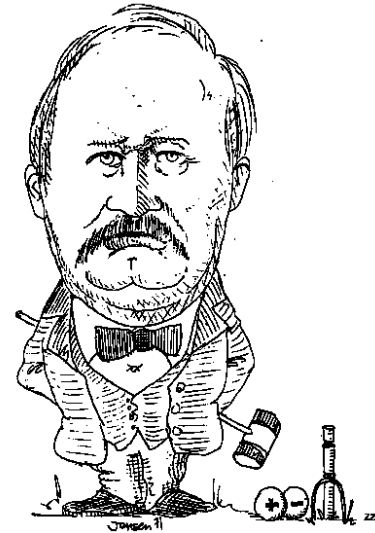
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$$k = zp e^{-E_a/RT}$$

$$k = A e^{-E_a/RT}$$



*'pre exponential' factor*  
(slight dependence on  $T$   
will be ignored)



Svante August Arrhenius (19 February 1859 – 2 October 1927) was a Swedish scientist, originally a physicist, but often referred to as a chemist, and one of the founders of the science of physical chemistry.

temperature dependence of rate constant (**HW 10 #69 Z15.73**)

---

$$k_{T_2} = A e^{-E_a/RT_2}$$

$$k_{T_1} = A e^{-E_a/RT_1}$$

$$\frac{k_{T_2}}{k_{T_1}} = \frac{e^{-E_a/RT_2}}{e^{-E_a/RT_1}} = e^{-\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$

$$k_{T_2} = k_{T_1} e^{-\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$

$$\ln k_{T_2} = \ln k_{T_1} + \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \quad \text{eqn 15.11}$$

know  $k_{T_1}$  and  $E_a \Rightarrow$  get  $k_{T_2}$  or measure  $k_{T_1}$  and  $k_{T_2} \Rightarrow$  get  $E_a$

**Chemistry 1B**  
**SAMPLE FINAL EXAMINATION QUESTIONS**

$h = 6.626 \times 10^{-34} \text{ J s}$   
 $c = 3.000 \times 10^8 \text{ m/s}$   
 $m_e = 9.109 \times 10^{-31} \text{ kg}$   
 $m_p = 1.672 \times 10^{-27} \text{ kg}$   
 $R = 8.3145 \text{ JK}^{-1}\text{mol}^{-1}$   
 $R = 0.08206 \text{ L atm K}^{-1}\text{mol}^{-1}$

$$\tilde{\nu} = \frac{1}{\lambda} = \frac{\nu}{c}$$

$$\frac{1}{\lambda} = Z^2 (1.097 \times 10^7 \text{ m}^{-1}) \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \quad n_2 > n_1$$

$$(m\Delta u) \Delta x \geq \frac{h}{4\pi}$$

$$E \approx -2.18 \times 10^{-18} \text{ J} \frac{Z_{\text{eff}}^2}{n^2}$$

$$r \approx (5.29 \times 10^{-12} \text{ m}) \frac{n^2}{Z_{\text{eff}}}$$

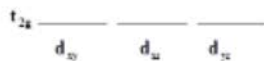
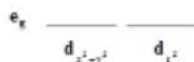
wavelength range of visible light :  
 $\lambda \approx 3.5 \times 10^{-7} \text{ m}$  to  $7.0 \times 10^{-7} \text{ m}$

$$k = Ae^{-E_a/RT}$$

$$\ln k = \ln A - \frac{E_a}{RT}$$

$$t_{1/2} = \frac{0.693}{k} \quad \text{for first-order reaction}$$

Energy levels of d-orbitals in octahedral crystal field:



Note: SAMPLE FINAL includes questions that are **representative** of the material, but do not include all possible topics.

| bond | approximate vibrational group frequency (cm <sup>-1</sup> ) |
|------|---|
| C-C  | ~ 1000-1400   |
| C=C  | ~1600   |
| C-O  | ~ 1100  |
| C=O  | ~ 1800  |
| C-N  | ~ 1000  |
| C≡N  | ~ 2100-2200   |
| C-H  | ~ 2800-3200   |
| N-H  | ~ 3300 (weak)<br>~1600 (intense)                            |
| O-H  | ~ 3600  |
| C-Cl | ~ 550-800   |

**Integrated Rate Laws:**

$$[A] = -kt + [A]_0 \quad \text{zeroth-order}$$

$$\ln[A] = -kt + \ln[A]_0 \quad \text{first-order}$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0} \quad \text{second-order}$$

## Example (by request)

---

A first-order reaction with activation energy  $E_a=50 \text{ kJ mol}^{-1}$  has a rate constant of  $k_{300}=3.0 \times 10^{-1} \text{ sec}^{-1}$  at 300K.

What is the rate constant at 310K?

What is the ratio of  $k_{310}/k_{300}$  ?



*adage raise 10K → twice as fast*

---

# Arrhenius equation

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From Wikipedia, the free encyclopedia

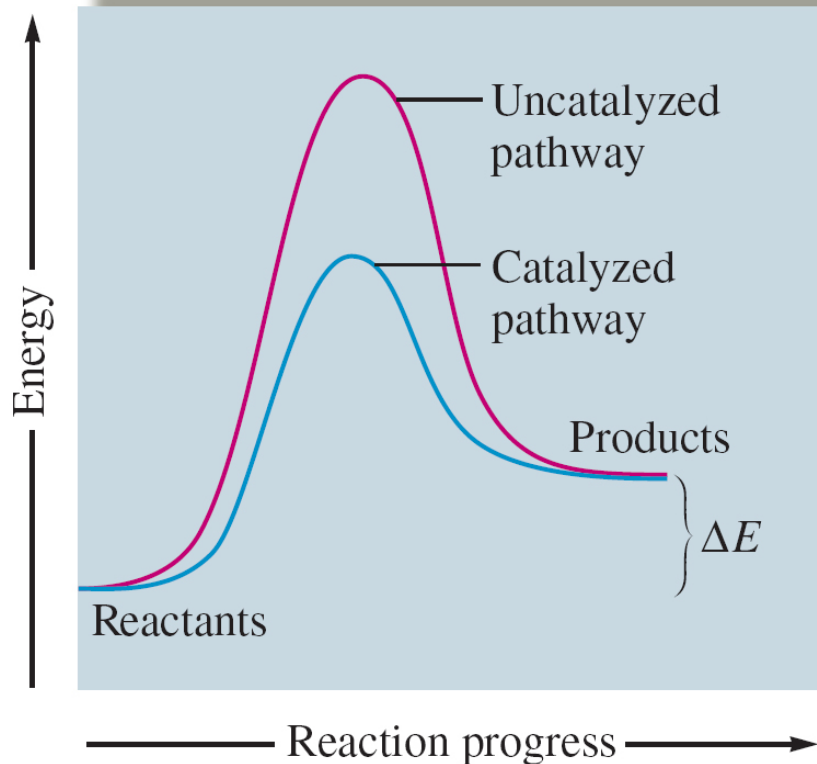
The **Arrhenius equation** is a simple, but remarkably accurate, formula for the temperature dependence of the **rate constant**, and therefore, rate of a chemical reaction.<sup>[1]</sup> The equation was first proposed by the Dutch chemist **J. H. van 't Hoff** in 1884; five years later in 1889, the Swedish chemist **Svante Arrhenius** provided a physical justification and interpretation for it. Nowadays it is best seen as an **empirical** relationship.<sup>[2]</sup> It can be used to model the temperature-variance of diffusion coefficients, population of crystal vacancies, creep rates, and many other thermally-induced processes/reactions.

A historically useful generalization supported by the Arrhenius equation is that, for many common chemical reactions at room temperature, the reaction rate doubles for every 10 degree Celsius increase in temperature.

$$E_a \approx 50 \text{ kJ mol}^{-1} \quad \text{at } T=298 \text{ K}$$

## catalyst ?

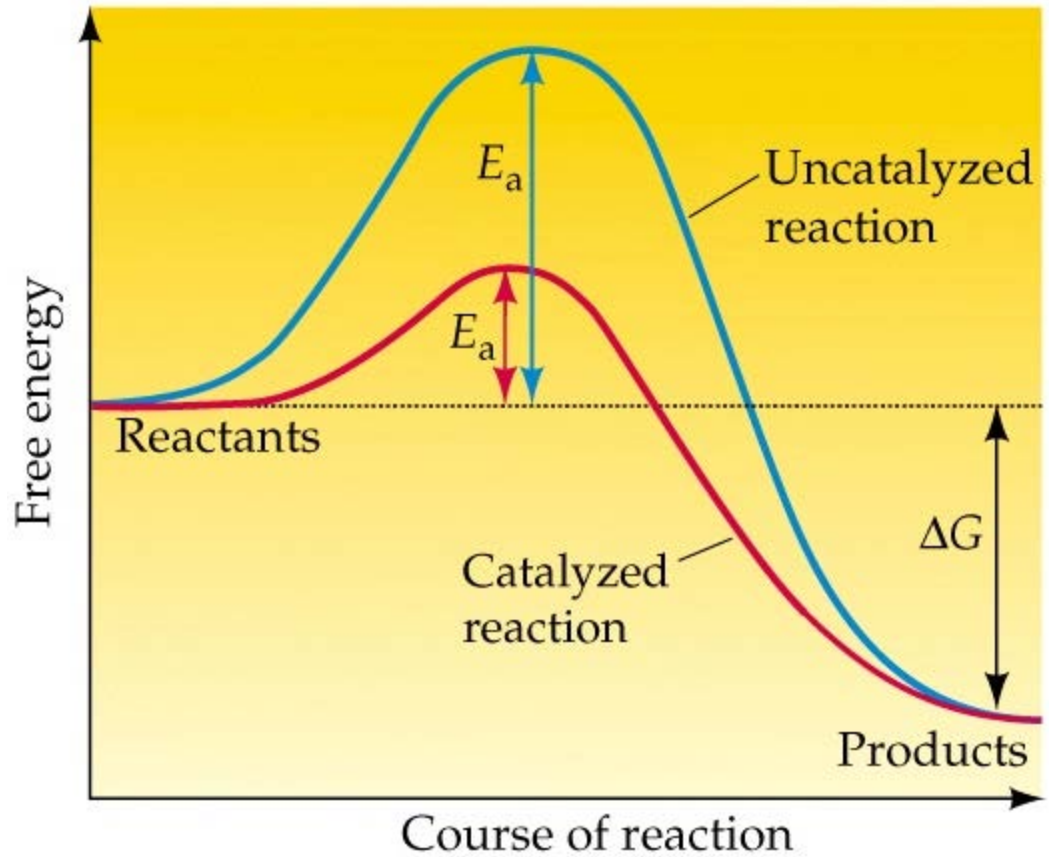
substance that increases the rate of a chemical reaction without being consumed itself



- catalysis changes the reaction pathway
- catalyzed pathway has lower  $E_a$  but same  $\Delta E = E_{\text{products}} - E_{\text{reactants}}$



## Catalysis and activation energy

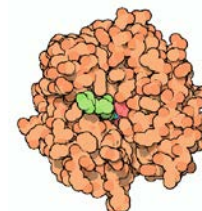
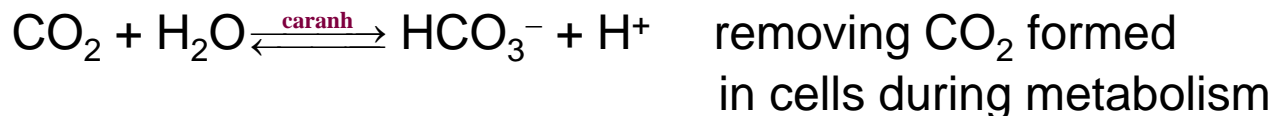


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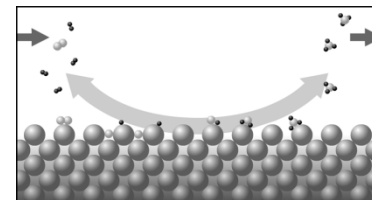
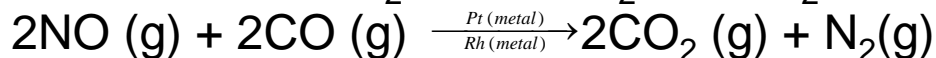
<http://www.columbia.edu/cu/biology/courses/c2005/purves6/figure06-14.jpg>

# general types of catalysis

- enzymatic catalysis  
e.g. carbonic anhydrase catalyzes



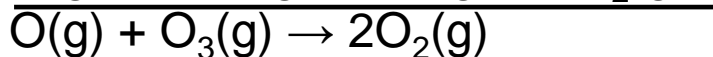
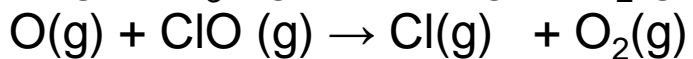
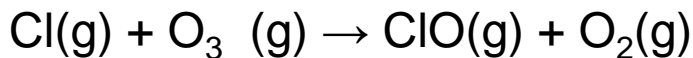
- surface (heterogeneous) catalysis  
e.g. platinum and rhodium metallic particles in catalytic converter  
converts CO, NO, N<sub>2</sub>O and NO<sub>2</sub> to CO<sub>2</sub> and N<sub>2</sub>



- homogeneous catalysis (ozone depletion)  
e.g.  $\text{O}(\text{g}) + \text{O}_3(\text{g}) \rightarrow 2\text{O}_2(\text{g})$

**E<sub>a</sub>=17.1 kJ mol<sup>-1</sup> uncatalyzed**

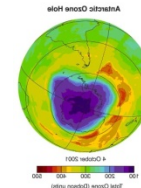
*catalyzed by Cl<sup>-</sup>*

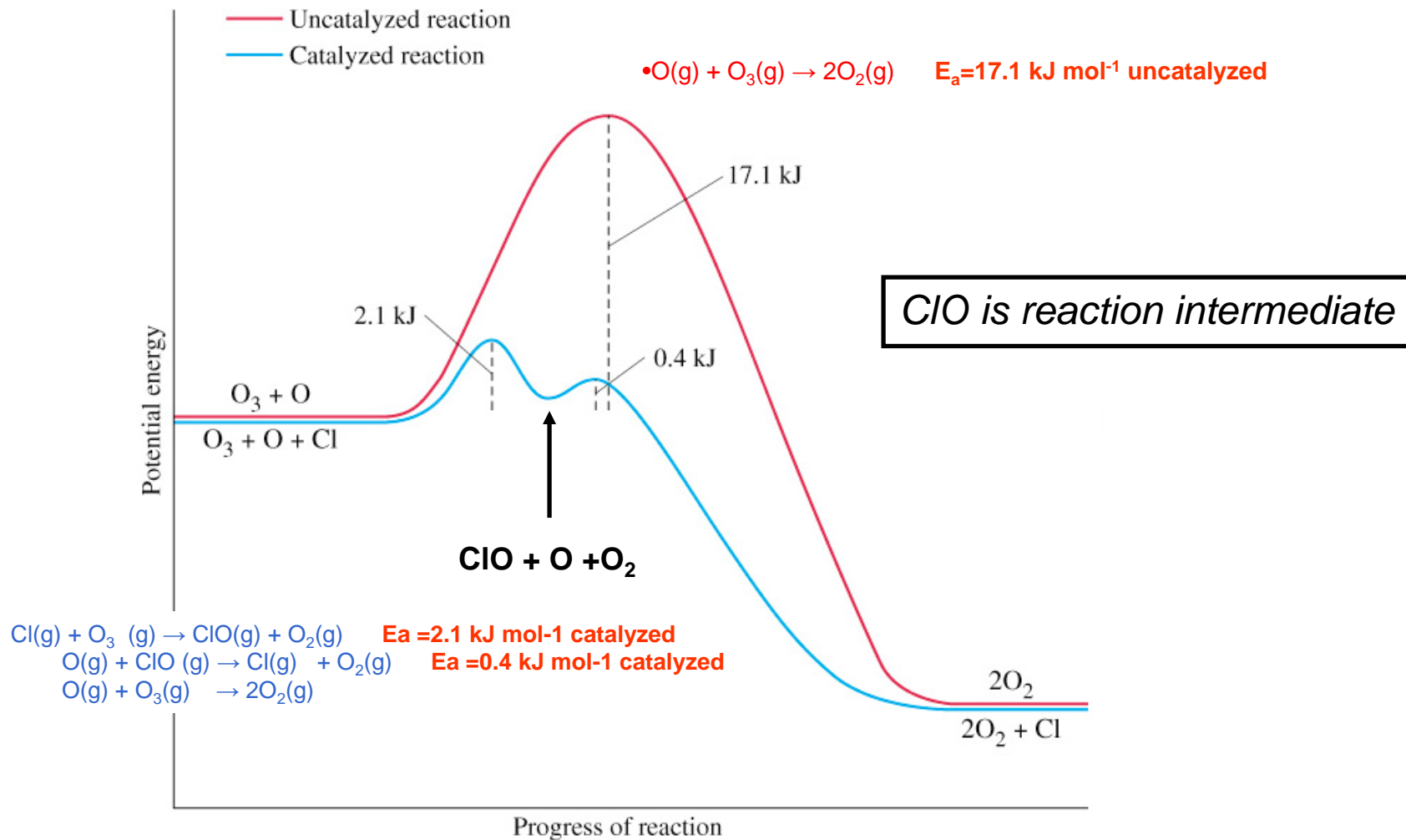
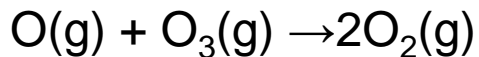


*from CCl<sub>2</sub>F<sub>2</sub> + hν → CClF<sub>2</sub> + Cl<sup>-</sup>*

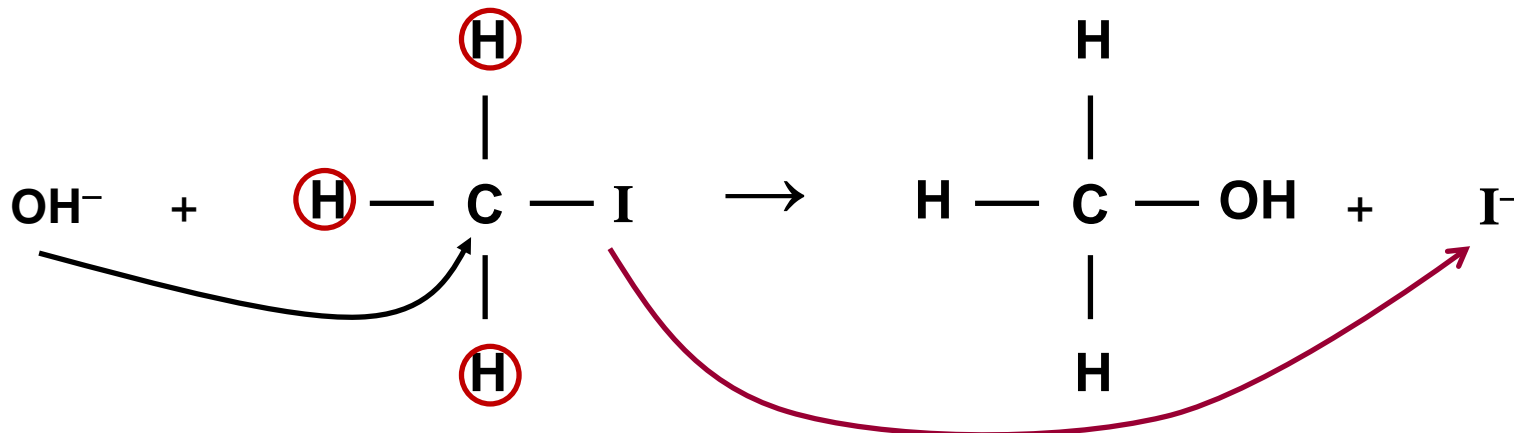
**E<sub>a</sub> =2.1 kJ mol<sup>-1</sup> catalyzed**

**E<sub>a</sub> =0.4 kJ mol<sup>-1</sup> catalyzed**





one last example of mechanism:  $\text{OH}^- + \text{H}_3\text{CI} \rightarrow \text{CH}_3\text{OH} + \text{I}^-$  vs  $\text{OH}^- + (\text{CH}_3)_3\text{CI} \rightarrow (\text{CH}_3)_3\text{COH} + \text{I}^-$

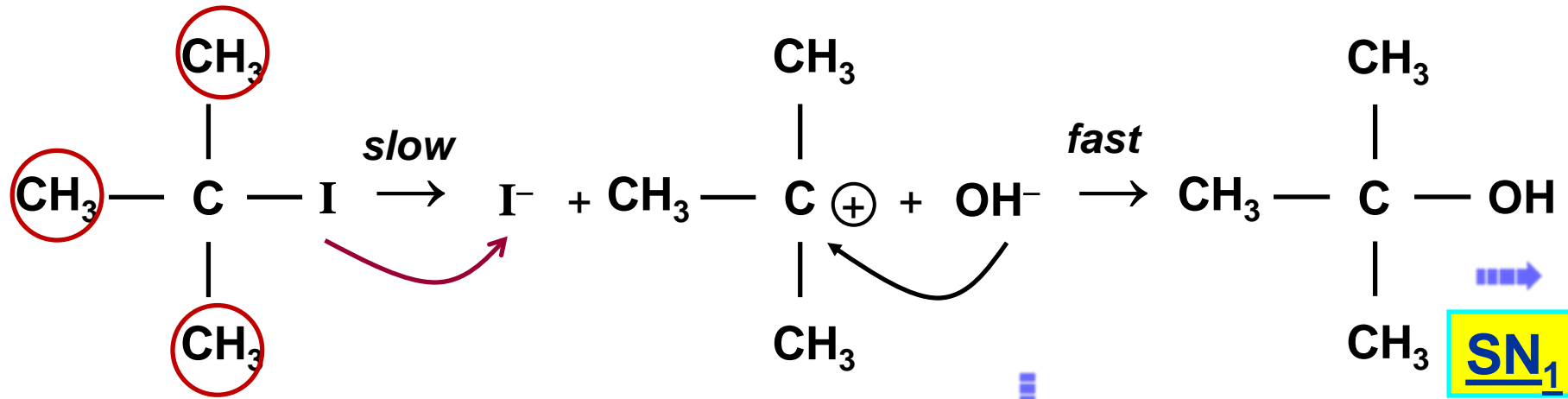


**SN<sub>2</sub>**

reaction rate =  $k_{\text{SN2}} [\text{OH}^-] [\text{CH}_3\text{CI}]$

second-order

bimolecular

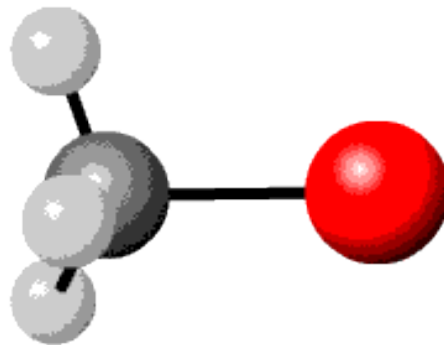
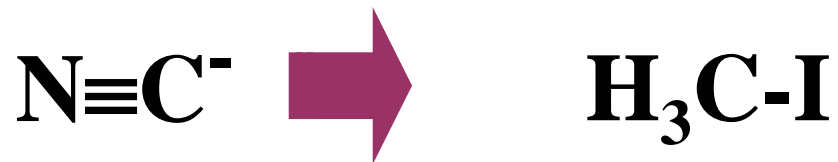


**SN<sub>1</sub>**

reaction rate =  $k_{\text{SN1}} [\text{C}_4\text{H}_9\text{I}]$   
 no  $[\text{OH}^-] \Rightarrow [\text{OH}^-]^0$

slow followed by fast

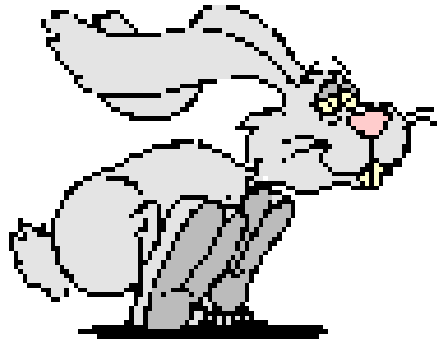
first-order



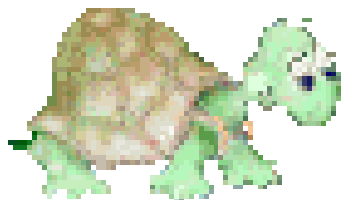
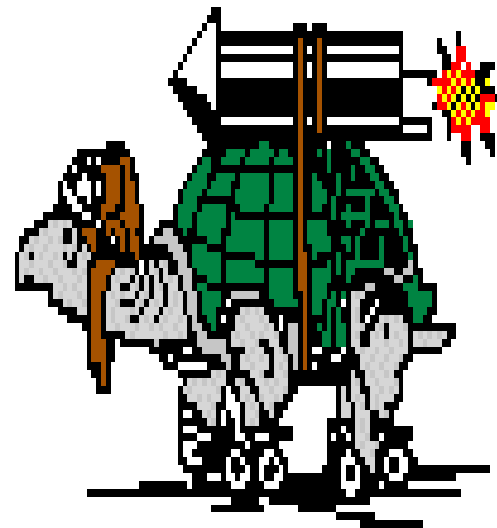
**Kinetics: How fast a reaction proceeds, and the molecular steps involved in a reaction [*the mechanism of a reaction*].**

*kinetic measurements: how fast will a reaction proceed*

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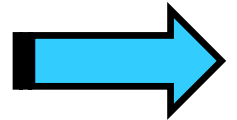
Catalysis



- ☑ • kinetics and mechanism of chemical reaction
- ☑ • energy profile and reaction coordinate
- ☑ • activation energy and temperature dependence of rate constant
- ☑ • catalysis

***FINIS !!!***

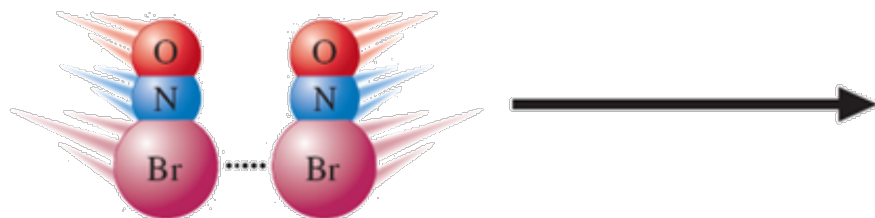
***BUT ONE MORE THING***





*p* the orientation factor ( $2 \text{ ONBr} \rightarrow 2 \text{ NO} + \text{Br}_2$ )

what relative orientation of two ONBr molecules will lead to a reaction ?  
(Br-----Br bond formation)



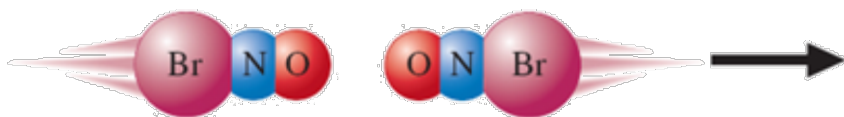
yes

(a)



yes

(b)



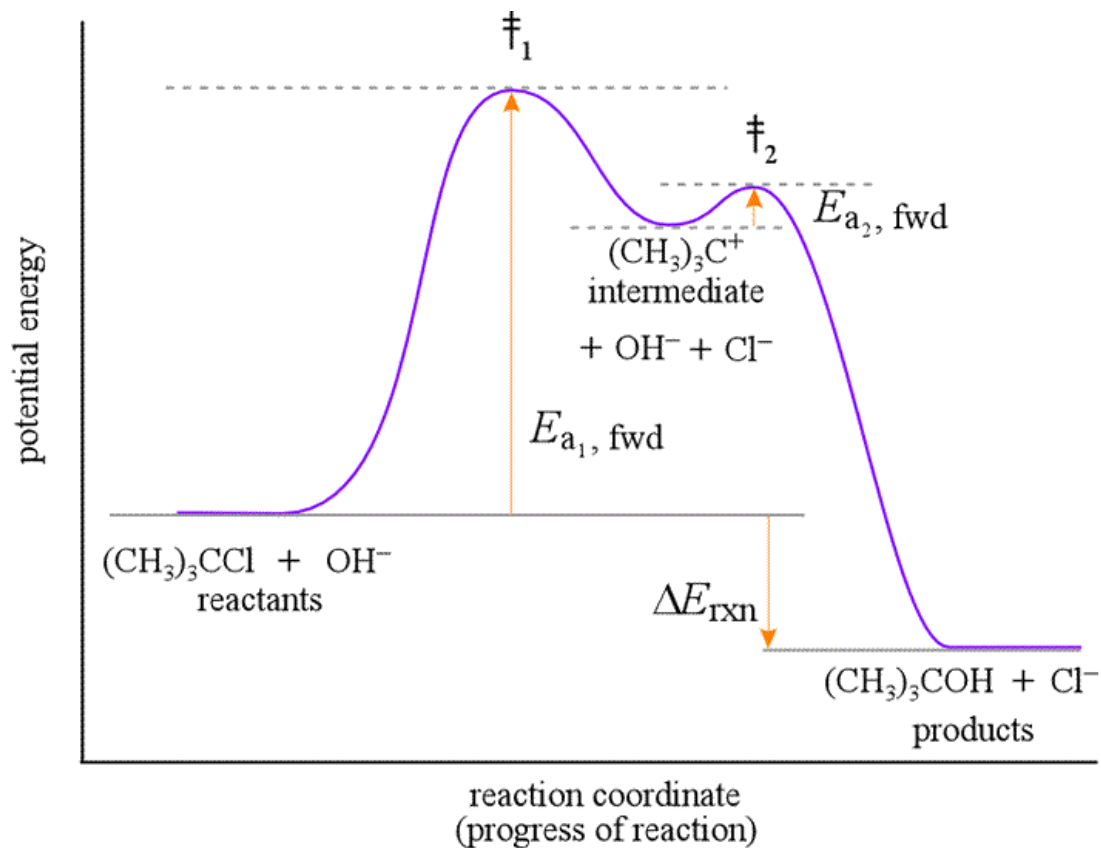
no

(c)

the more 'particular' the reaction is about alignment of colliding molecule the smaller the *p* factor



two-step, slow then fast, first-order with reaction intermediate  
HW#10 probs 67 and 68



[http://guweb2.gonzaga.edu/faculty/cronk/CHEM240/images/reaction\\_coordinate\\_SN1.gif](http://guweb2.gonzaga.edu/faculty/cronk/CHEM240/images/reaction_coordinate_SN1.gif)



## Example (by request)

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A first-order reaction with activation energy  $E_a=50 \text{ kJ mol}^{-1}$  has a rate constant of  $k_{300}=3.0 \times 10^{-1} \text{ sec}^{-1}$  at 300K.

What is the rate constant at 310K?

What is the ratio of  $k_{310}/k_{300}$  ?

$$\ln k_T = \ln A - \frac{E_a}{RT}$$

$$1. \quad \ln k_{310} = \ln A - \frac{E_a}{R(310K)}$$

$$2. \quad \ln k_{300} = \ln A - \frac{E_a}{R(300K)}$$

(same A and  $E_a$ )

subtract: 1 - 2

$$\begin{aligned} \ln \left( \frac{k_{310}}{k_{300}} \right) &= -\frac{E_a}{R} \left( \frac{1}{310K} - \frac{1}{300K} \right) \\ &= -\frac{50 \times 10^3 \text{ J mol}^{-1}}{8.3145 \text{ J mol}^{-1} \text{ K}^{-1}} \left( \frac{1}{310K} - \frac{1}{300K} \right) \quad \text{note unitless} \end{aligned}$$

## example continued

---

$$\ln \left( \frac{k_{310}}{k_{300}} \right) = -\frac{E_a}{R} \left( \frac{1}{310\text{K}} - \frac{1}{300\text{K}} \right)$$
$$= -\frac{50 \times 10^3 \text{ J mol}^{-1}}{8.3145 \text{ J mol}^{-1} \text{ K}^{-1}} \left( \frac{1}{310\text{K}} - \frac{1}{300\text{K}} \right) \quad \text{note unitless}$$

$$\ln \left( \frac{k_{310}}{k_{300}} \right) = 0.6466$$

$$k_{310} = k_{300} \times e^{0.6466} = (3.0 \times 10^{-1} \text{ sec}^{-1}) \times 1.91 = 0.574 \text{ sec}^{-1}$$

$$\ln \left( \frac{k_{310}}{k_{300}} \right) = 0.6466 \Rightarrow \frac{k_{310}}{k_{300}} = 1.91 \approx 2$$



*Old adage:*

raise T by  $10^\circ$  (300  $\rightarrow$  310)  
double reaction rate  
( $E_a \approx 50 \text{ kJ/mol}$ )

